

## **Chapter 6 - Electrolysis and its uses**

### **Electrolysis of aluminium oxide**

Anodising

### **Extraction of sodium**

### **Electrolysis of aqueous solutions**

Electrolysis of water

Chlor-alkali industry

### **Purification of copper**

### **Electrolysis guidelines**

### **Electroplating**

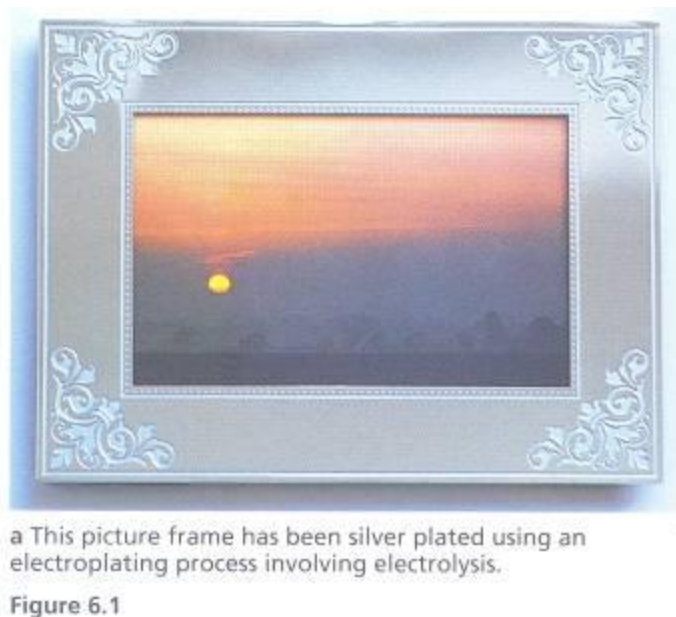
Plating plastics

### **Calculations in electrolysis**

### **Checklist**

### **Additional questions**

What do all the items in the photographs shown in Figure 6.1 have in common? They all involve electricity through a process known as **electrolysis**. Electrolysis is the process of splitting up (decomposing) substances by passing an electric current through them. The substance which is decomposed is called the **electrolyte** (Figure 6.2). An electrolyte is a substance that conducts electricity when in the molten state or in solution. The electricity is carried through the electrolyte by **ions**. The electric current enters and leaves the electrolyte through **electrodes**, which are usually made of unreactive metals such as platinum or of the non-metal carbon (**inert** electrodes). The names given to the two electrodes are **cathode**, the negative electrode which attracts **cations** (positively charged ions), and **anode**, the positive electrode which attracts **anions** (negatively charged ions).



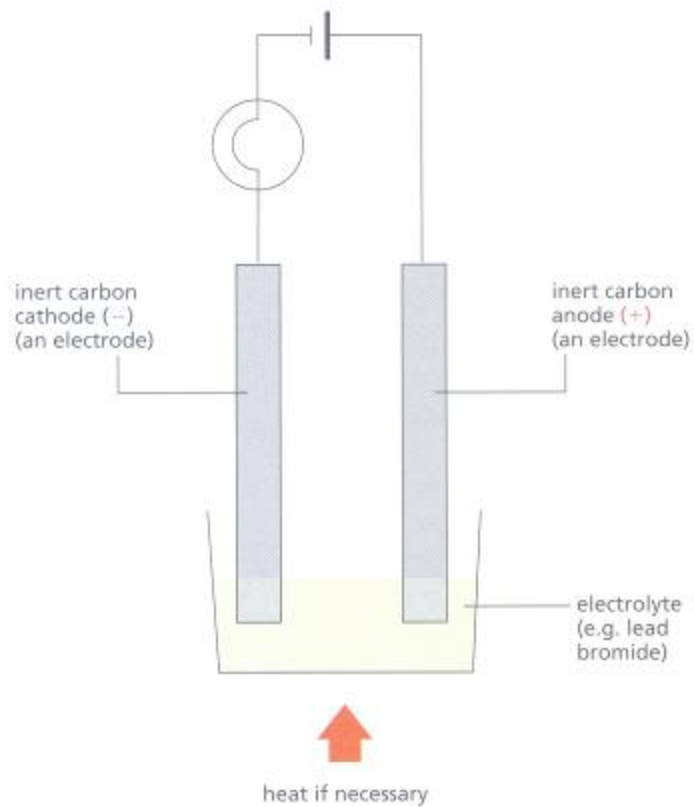
Electrolysis is very important in industry. To help you to understand what is happening in the process shown in the photographs, we will first consider the electrolysis of aluminium oxide.



**b** Aluminium is produced by electrolysis.



**c** This watch has a thin coating of gold over steel; the thin coating is produced by electrolysis.



**Figure 6.2** The important terms used in electrolysis.

## **Electrolysis of aluminium oxide**

Aluminium is the most abundant metallic element in the Earth's crust. It was discovered in 1825 by Hans Christian Oersted in Copenhagen, Denmark, but was first isolated by Friedrich Wöhler in 1827. It makes up 8% of the crust and is found in the minerals bauxite (Figure 6.3), cryolite and mica, as well as in clay.

In the nineteenth century Napoleon III owned a very precious dinner service. It was said to be made of a metal more precious than gold. That metal was aluminium. The reason it was precious was that it was very rarely found as the pure metal. Aluminium is a reactive metal and as such was very difficult to extract from its ore. Reactive metals hold on tightly to the element(s) they have combined with and many are extracted from their ores by electrolysis.

Today we use aluminium in very large quantities. The annual production in the world is 19.5 million tonnes, of which 248 000 tonnes are produced in the UK. The commercial extraction of aluminium has been made possible by two scientists, working independently of each other, who discovered a method using electrolysis. The two scientists were Charles Martin Hall (USA), who discovered the process in 1886, and the French chemist Paul Héroult, who discovered the process independently in the same year.

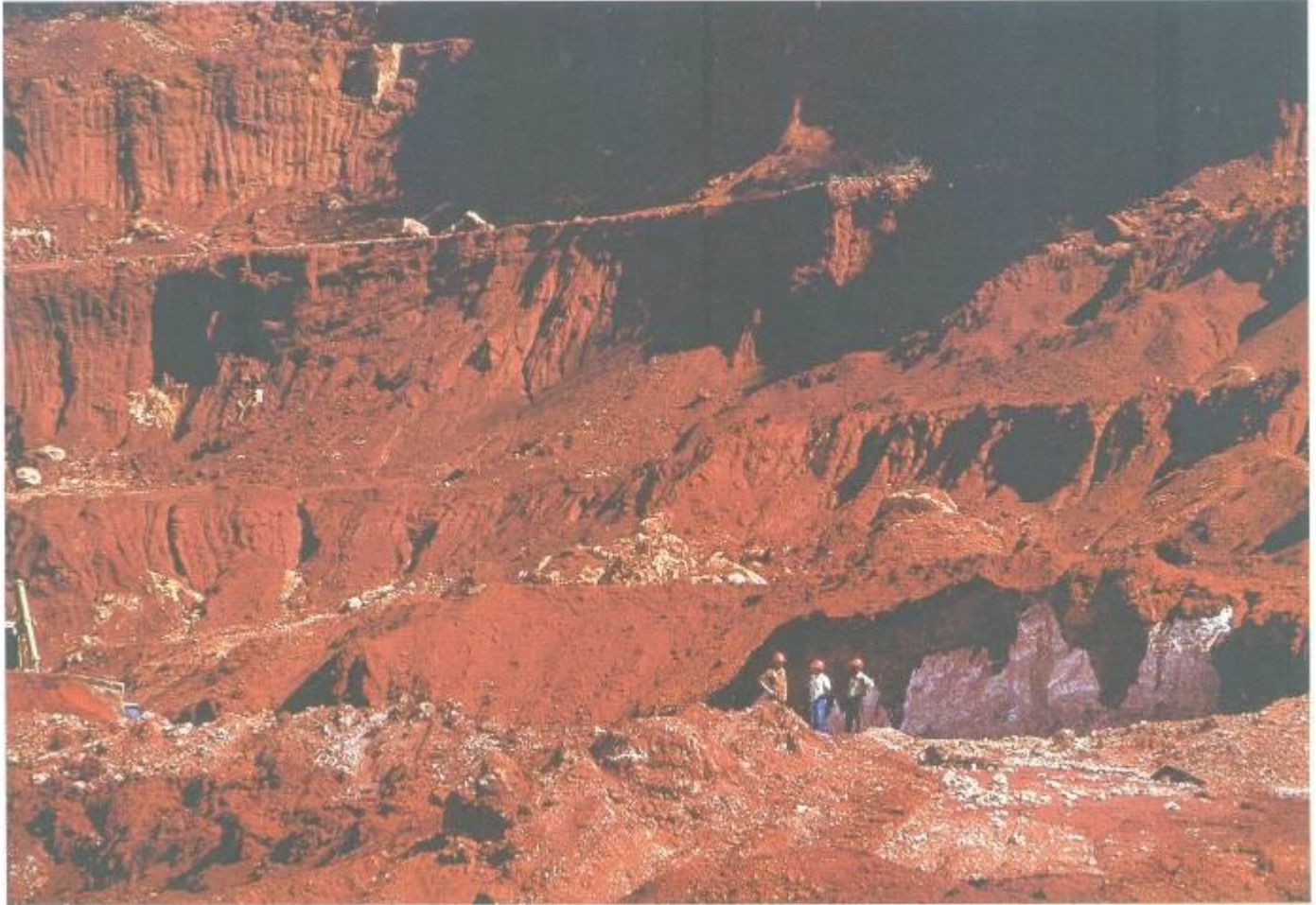


Figure 6.3 Bauxite mining in Australia.

The process they developed, often called the Hall–Héroult process, involves the electrolysis of aluminium oxide (alumina). The process involves the following stages.

- Bauxite, an impure form of aluminium oxide, is first treated with sodium hydroxide to obtain pure aluminium oxide, removing impurities such as iron (III) oxide and sand.
- The purified aluminium oxide is then dissolved in molten cryolite ( $\text{Na}_3\text{AlF}_6$ ). Cryolite, a mineral found naturally in Greenland, is used to reduce the working temperature of the Hall–Héroult cell from  $2017^\circ\text{C}$  (the melting point of pure aluminium oxide) to between  $800$  and  $1000^\circ\text{C}$ . Therefore, the cryolite provides a considerable saving in the energy requirements of the process. In recent years it has become necessary to manufacture the cryolite.



- The molten mixture is then electrolysed in a cell similar to that shown in Figure 6.4.

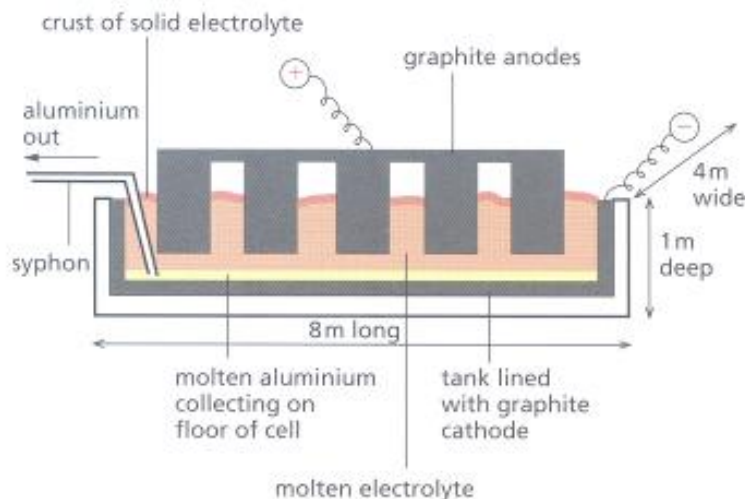
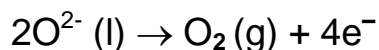


Figure 6.4 The Hall-Héroult cell is used in industry to extract aluminium by electrolysis.

The anodes of this process are blocks of graphite which are lowered into the molten mixture from above. The cathode is the graphite lining of the steel vessel containing the cell.

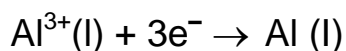
Aluminium oxide is an ionic compound. When it is melted the ions become mobile, as the strong electrostatic forces of attraction between them are broken by the input of heat energy. During electrolysis the negatively charged oxide ions are attracted to the anode (the positive electrode), where they lose electrons to form oxygen gas.

oxide ions  $\rightarrow$  oxygen molecules + electrons



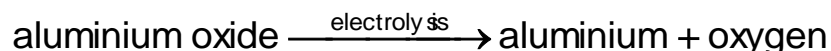
The positive aluminium ions are attracted to the cathode (the negative electrode). They gain electrons to form molten aluminium metal.

aluminium ions + electrons  $\rightarrow$  aluminium metal

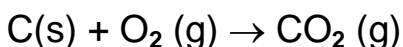
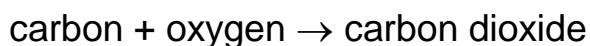


The gain of electrons which takes place at the cathode is called **reduction**. The loss of electrons which takes place at the anode is called **oxidation**. This is a further way of considering oxidation and reduction (Chapter 2, p. 17). A handy way of remembering it is **oil rig** (oxidation is loss, reduction is gain of electrons).

The overall reaction which takes place in the cell is:



The molten aluminium collects at the bottom of the cell and it is syphoned out at regular intervals. No problems arise with other metals being deposited, since the cryolite is largely 'unaffected' by the flow of electricity. Problems do arise, however, with the graphite anodes. At the working temperature of the cell, the oxygen liberated reacts with the graphite anodes, producing carbon dioxide.



The anodes burn away and have to be replaced on a regular basis.

The electrolysis of aluminium oxide is a continuous process in which vast amounts of electricity are used. Approximately 15 kWh of electricity are used to produce 1 kg of aluminium. In order to make the process an economic one, a cheap form of electricity is required. Hydroelectric power (HEP) is usually used for this process. The plant shown in Figure 6.5 uses an HEP scheme to provide some of the electrical energy required for this process. Further details about HEP are given in Chapter 13.



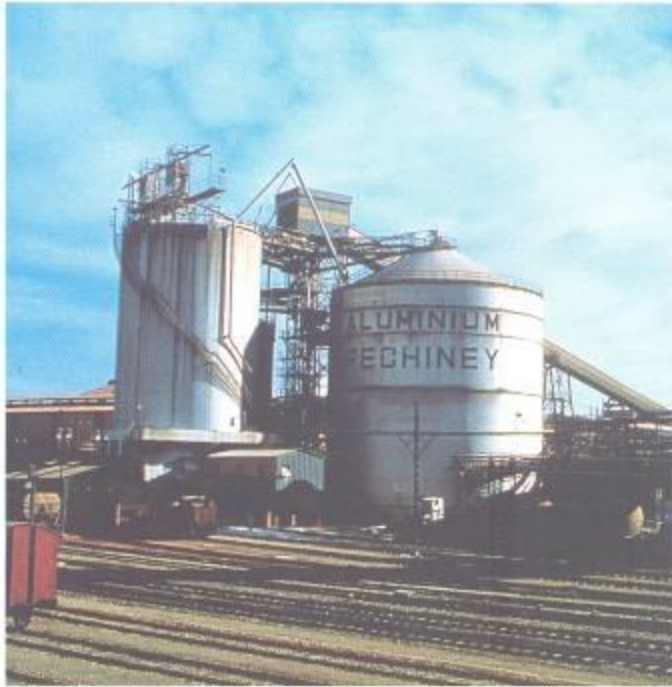


Figure 6.5 An aluminium smelting plant.

Using cheap electrical energy has allowed aluminium to be produced in such large quantities that it is the second most widely used metal after iron. It is used in the manufacture of electrical cables, cars, bikes and cooking foil as well as in alloys (Chapter 9) such as duralumin, which is used in the manufacture of aeroplane bodies (Figure 6.6).



Figure 6.6 Aluminium is used in the manufacture of aeroplane bodies.



Figure 6.7 Bauxite pollution of the Amazon in Brazil.

Environmental problems associated with the location of aluminium plants are concerned with:

- the effects of the extracted impurities, which form a red mud (Figure 6.7)
- the fine cryolite dust, which is emitted through very tall chimneys so as not to affect the surrounding area
- the claimed link between environmental aluminium and a degenerative brain disorder called Alzheimer's disease — it is thought that aluminium is a major influence on the early onset of this disease. However, the evidence is still inconclusive

## Anodising

This is a process in which the surface coating of oxide on aluminium ( $\text{Al}_2\text{O}_3$ ) is made thicker. In this process the aluminium object is made the anode in a cell in which the electrolyte is dilute sulphuric acid. During the



**Figure 6.8** The oxide layer on the surface of these rolls of aluminium has been thickened, and dyes added to obtain the vibrant colours.

the electrolysis process, oxygen is produced at the anode and combines with the aluminium. The oxide layer on the surface of the aluminium therefore increases. Dyes can be mixed with the electrolyte and so the new thicker coating of oxide is colourful and also decorative (Figure 6.8).

## Question

1. Produce a flow chart to summarise the processes involved in the extraction of aluminium metal from bauxite.

## Extraction of sodium

Sodium, a very reactive metal, was first discovered by Sir Humphry Davy in 1807. It is now extracted by electrolysing molten sodium chloride in the Down's cell. Sodium chloride may be obtained from the evaporation of sea water or mined as rock salt (Figure 6.9). The largest salt mines in England are those in Cheshire. The rock salt is mined from large under-ground caverns. Rock salt for the chemical industry is obtained by dissolving the salt and pumping the brine to the surface.



Figure 6.9 Rock salt mining in Argentina.

The first economic extraction of sodium was developed by Down in 1921 in North America using hydroelectric power generated by the Niagara Falls. This ensured the availability of cheap electricity.

The process involves the following.

- Sodium chloride, obtained from rock salt or sea water, is mixed with calcium chloride and melted. The calcium chloride is added to the sodium chloride electrolyte to reduce the working temperature of the cell from 801 °C (the melting point of sodium chloride) to 600°C (the melting point of the mixture). This saves electrical energy and, therefore, makes the process more economical.
- The mixture is then electrolysed in a cell similar to the one in Figure 6.10.

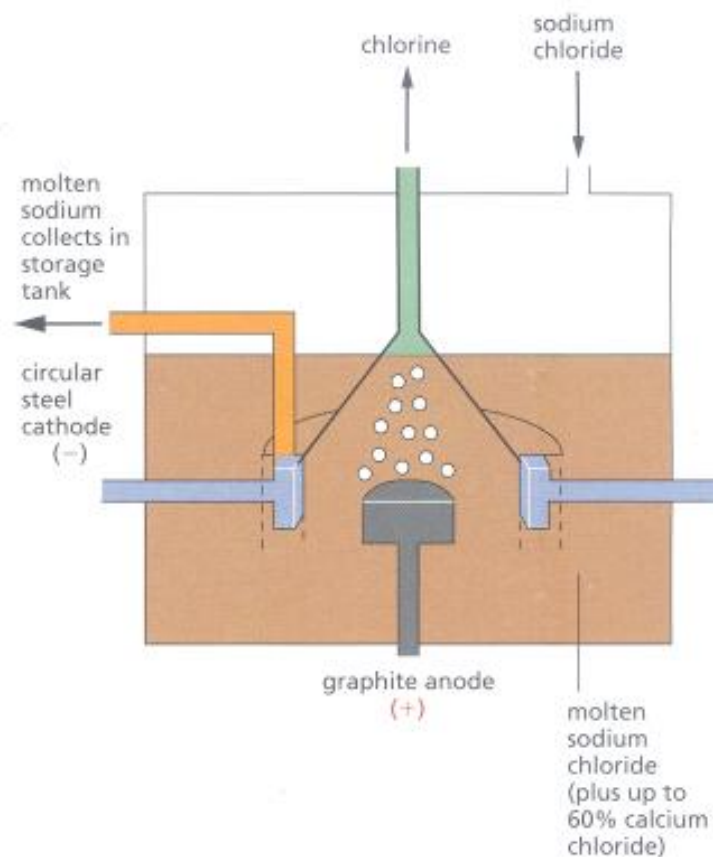


Figure 6.10 The Down's cell used for the electrolysis of sodium chloride.

The chloride ions are attracted to the anode, where they lose electrons and form chlorine gas.

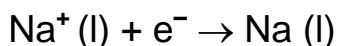




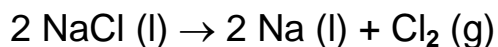
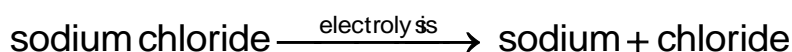


The positive sodium ions are attracted to the cathode. They gain electrons to form molten sodium metal.

sodium ions + electrons  $\rightarrow$  sodium metal



The overall reaction which takes place in the cell is:



The cathode is a circle of steel around the graphite anode. At 600°C sodium and chlorine would react violently together to reform sodium chloride. To prevent this from happening, the Down's cell contains a steel gauze around the graphite anode to keep it and the cathode apart. The molten sodium floats on the electrolyte and is run off for storage.

A problem arises, however, in that calcium ions are also attracted to the cathode, where they form calcium metal. Therefore, the sodium which is run off contains a significant proportion of calcium. Fortunately, the calcium crystallises out when the mixture cools and relatively pure sodium metal remains (Figure 6.11).



Figure 6.11 Molten sodium – the result of electrolysis of sodium chloride.

The worldwide production of sodium each year is 60 - 80 000 tonnes. Of this, 16 000 tonnes are produced per year in the UK.

Sodium is used as a liquid coolant in nuclear power stations as well as in street lighting and in the production of metals such as titanium. Chlorine gas (the co-product) is sold to make the process more economical.

## Questions

1. Use your research skills, including the Internet, to find uses, other than those given in the text, for sodium.
2. A student carries out the electrolysis of molten lead(II) bromide in a fume

cupboard.

- a. Draw a diagram to show a suitable apparatus the student could use to carry out this experiment.
- b. Write anode and cathode reactions to represent the processes taking place during the electrolysis.
- c. Why does this experiment need to be carried out in a fume cupboard?
- d. Find uses for the anode product of this cell.



## Electrolysis of aqueous solutions



Other industrial processes involve the electrolysis of aqueous solutions. To help you to understand what is happening in these processes, we will first consider the electrolysis of water.

### Electrolysis of water

Pure water is a very poor conductor of electricity because there are so few ions in it. However, it can be made to decompose if an electric current is passed through it in a Hofmann voltameter, as in Figure 6.12.

To enable water to conduct electricity better, some dilute sulphuric acid (or sodium hydroxide solution) is added. When the power is turned on and an electric current flows through this solution, gases can be seen to be produced at the two electrodes and they are collected in the side arms of the apparatus. After about 20 minutes, roughly twice as much gas is produced at the cathode as at the anode.

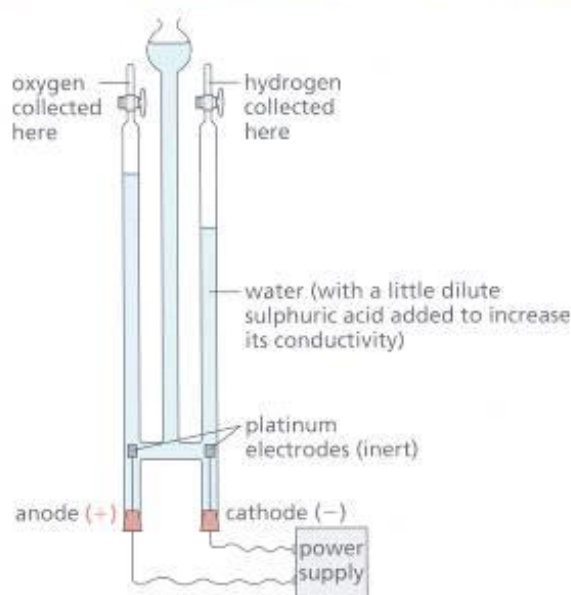


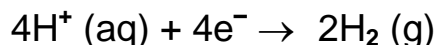
Figure 6.12 A Hofmann voltameter used to electrolyse water.

gas. For hydrogen to be collected in this way, the positively charged

The gas collected at the cathode burns with a squeaky pop, showing it to be hydrogen

hydrogen ions must have moved to the cathode.

hydrogen ions + electrons → hydrogen molecules



If during this process the water molecules lose  $\text{H}^+ (\text{aq})$ , then the remaining portion must be hydroxide ions,  $\text{OH}^- (\text{aq})$ . These ions are attracted to the anode. The gas collected at the anode relights a glowing splint, showing it to be oxygen.

This gas is produced in the following way.

hydroxide ions → water molecules + oxygen molecules + electrons



This experiment was first carried out by Sir Humphry Davy, who confirmed by this experiment that the formula for water was  $\text{H}_2\text{O}$ .

### **Chlor-alkali industry**

The electrolysis of saturated sodium chloride solution (brine) is the basis of a major industry. Rock salt (sodium chloride) is mined in the United Kingdom in Cheshire, Lancashire, Staffordshire and Cleveland. Three very important substances are produced in this electrolysis process – chlorine, sodium hydroxide and hydrogen. The electrolytic process is a very expensive one, requiring vast amounts of electricity. The process is economical only because all three products have a large number of uses (Figure 6.13).

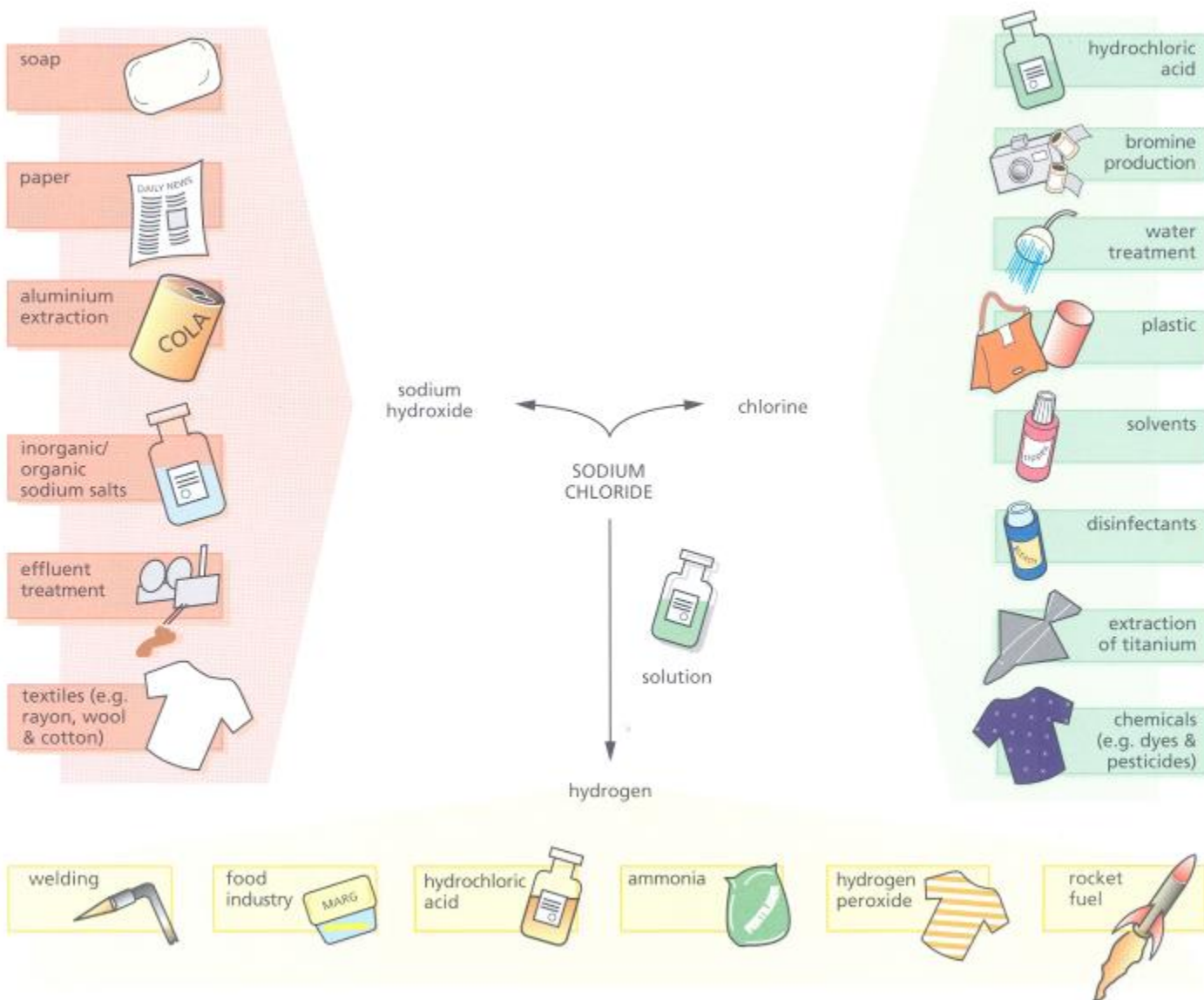


Figure 6.13 The chlor-alkali industry.

There are two well-established methods for electrolysing brine, the **diaphragm cell** and the **mercury cell**. However, recent developments in electrolysis technology, by chemical engineers, have produced the **membrane cell** (Figure 6.14). This method is now preferred to the other two because it produces a purer product, it causes less pollution and it is cheaper to run.

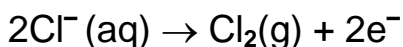
The brine is first purified to remove calcium, strontium and magnesium compounds by a process of ion exchange (see Chapter 8).

The membrane cell is used continuously, with fresh brine flowing into the cell as the process breaks up the brine. The cell has been designed to ensure that the products do not mix. The ions in this concentrated sodium chloride solution are:

from the water:	$\text{H}^+ (\text{aq})$	$\text{OH}^- (\text{aq})$
from the sodium chloride:	$\text{Na}^+ (\text{aq})$	$\text{Cl}^- (\text{aq})$

When the current flows, the chloride ions,  $\text{Cl}^- (\text{aq})$ , are attracted to the anode. Chlorine gas is produced by the electrode process.

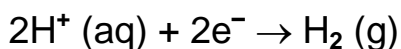
chloride ions  $\xrightarrow{\text{electrolysis}}$  chlorine molecules + electrons



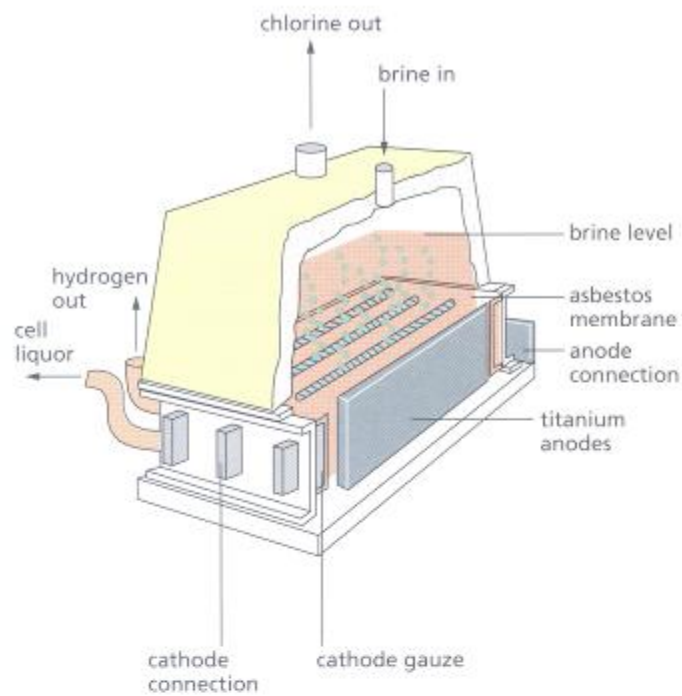
This leaves a high concentration of sodium ions,  $\text{Na}^+ (\text{aq})$ , around the anode.

The hydrogen ions,  $\text{H}^+(\text{aq})$ , are attracted to the cathode and hydrogen gas is produced.

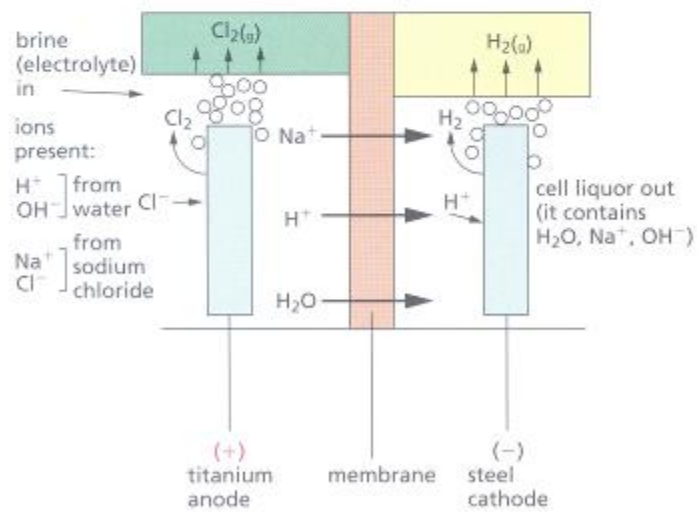
hydrogenions + electrons  $\xrightarrow{\text{reduction}}$  hydrogenmolecules



This leaves a high concentration of hydroxide ions,  $\text{OH}^- (\text{aq})$ , around the cathode. The sodium ions,  $\text{Na}^+ (\text{aq})$ , are drawn through the membrane, where they combine with the  $\text{OH}^- (\text{aq})$  to form sodium hydroxide, NaOH, solution. The annual production in the UK is 1.2 million tonnes and worldwide is 41 million tonnes.



a A section through the membrane cell.



b A diagrammatic representation of the reactions going on inside the cell.

Figure 6.14

## Questions

1. Suggest a reason for only 'roughly' twice as much hydrogen gas being produced at the cathode as oxygen gas at the anode in the electrolysis of water.

2. Account for the following observations which were made when concentrated sodium chloride solution, to which a little universal indicator had been added, was electrolysed in the laboratory in a Hofmann voltameter.

a. The universal indicator initially turns red in the region of the anode, but as the electrolysis proceeds it loses its colour.

b. The universal indicator turns blue in the region of the cathode.

3. Why is it important to remove compounds of calcium, strontium and magnesium before brine is electrolysed?

4. The uses of sodium hydroxide can be separated on a percentage basis as follows:

Neutralisation	5%
Paper manufacture	5%
Oil refining	5%
Soap/detergents	5%
Manufacture of rayon and acetate fibres	16%
Manufacture of chemicals	30%
Miscellaneous uses	34%

Use a graph plot program to create a 'pie' chart of this data.

## Purification of copper

Because copper is a very good conductor of electricity, it is used for electrical wiring and cables (Figure 6.15).

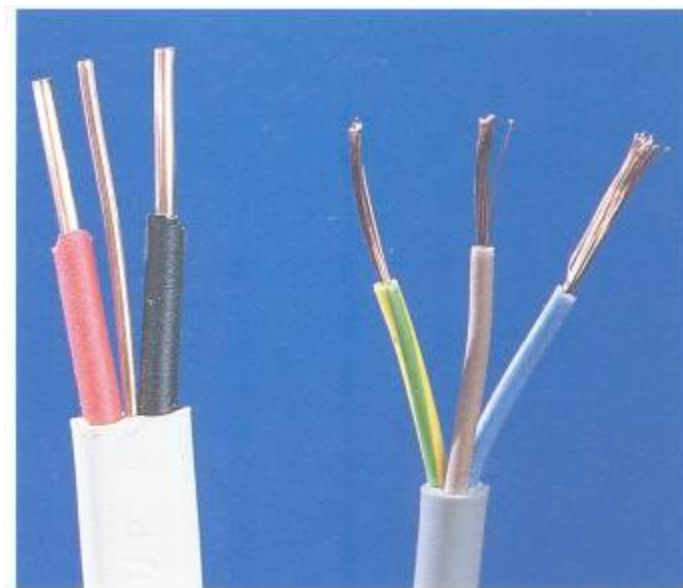


Figure 6.15 The copper used in electrical wiring has to be very pure.

However, even small amounts of impurities cut down this conductivity quite noticeably. The metal must be 99.99% pure to be used in this way. To ensure this level of purity, the newly extracted copper has to be purified by electrolysis (Figure 6.16).

The impure copper is used as the anode and is typically 1m square, 35-50 mm thick and 330 kg in weight. The cathode is a 1 mm thick sheet and weighs about 5 kg; it is made from very pure copper. The electrolyte is a solution of copper (II) sulphate ( $0.3 \text{ mol}\cdot\text{dm}^{-3}$ ) acidified with a  $2\text{mol}\cdot\text{dm}^{-3}$  solution of sulphuric acid to help the solution conduct electricity.

When the current flows, the copper moves from the impure anode to the pure cathode. Any impurities fall to the bottom of the cell and collect below the anode in the form of a slime. This slime is rich in precious metals and the recovery of these metals is an important aspect of the economics of the process. The electrolysis proceeds for about three weeks until the anodes are reduced to about 10% of their original size and the cathodes weigh between 100 and 120 kg. A potential of 0.25 V and a current density of 200



$\text{A}\cdot\text{m}^{-2}$  are usually used.

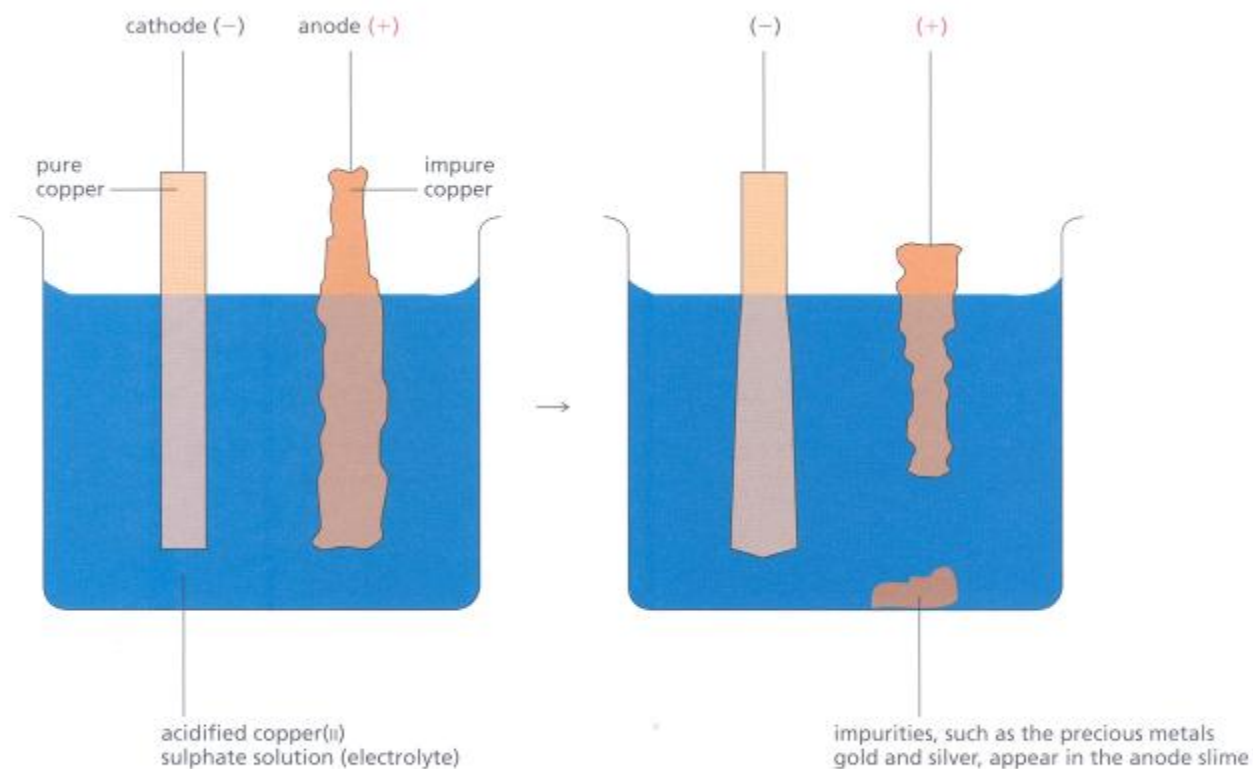


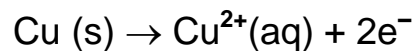
Figure 6.16 Copper purification process.

The ions present in the solution are:

from the water:	$\text{H}^+ (\text{aq})$	$\text{OH}^- (\text{aq})$
from the copper (II) sulphate:	$\text{Cu}^{2+} (\text{aq})$	$\text{SO}_4^{2-} (\text{aq})$

During the process the impure anode loses mass because the copper atoms lose electrons and become copper ions,  $\text{Cu}^{2+} (\text{aq})$  (Figure 6.17).

copper atoms  $\rightarrow$  copper ions + electrons



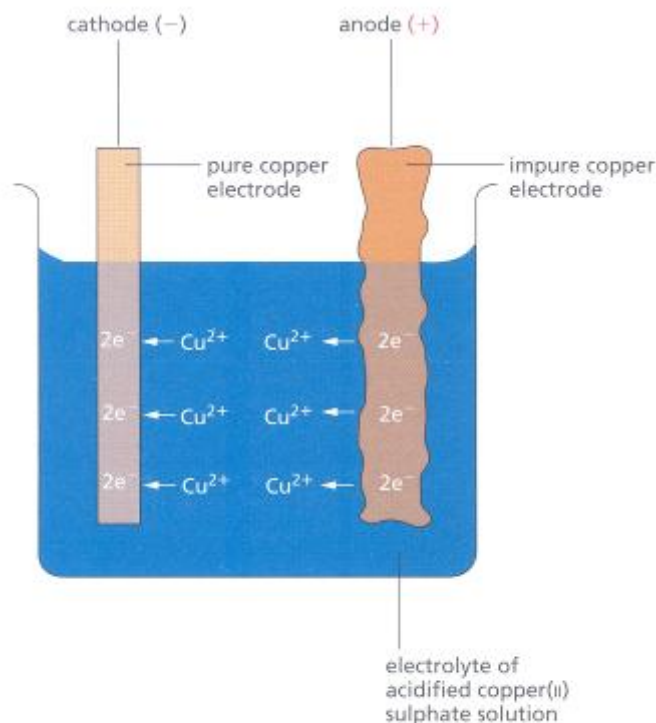
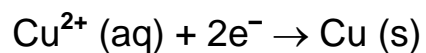


Figure 6.17 The movement of ions in the purification of copper by electrolysis.

The electrons released at the anode travel around the external circuit to the cathode. There the electrons are passed on to the copper ions,  $\text{Cu}^{2+} (\text{aq})$ , from the copper (II) sulphate solution and the copper is deposited or copper plated on to the cathode.

copper ions + electrons  $\rightarrow$  copper atoms



The annual production of copper worldwide is 12 million tonnes. We produce 57 000 tonnes in the UK, the vast majority of which is recycled.

## Electrolysis guidelines

The following points may help you work out the products of electrolysis in unfamiliar situations. They will also help you to remember what happens at each of the electrodes.

- Non-metals are produced at the anode whereas metals and hydrogen gas are produced at the cathode.
- At the anode, chlorine, bromine and iodine (the halogens) are produced in preference to oxygen.
- At the cathode, hydrogen is produced in preference to metals unless unreactive metals such as copper and nickel are present.

## Questions

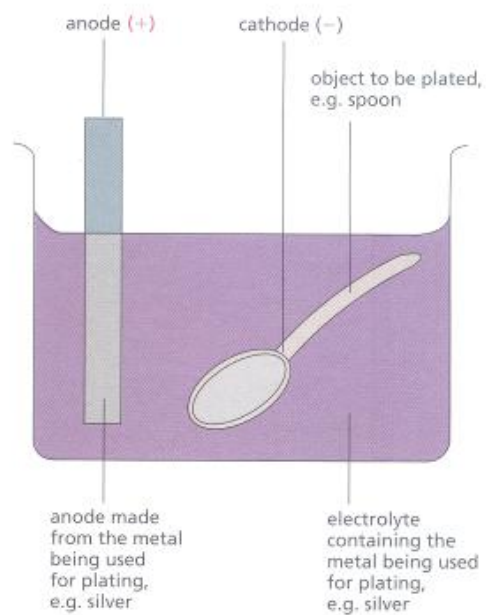
1. Why do you think it is advantageous to use inert electrodes in the electrolysis processes?
2. Predict the products of electrolysis of a solution of copper (II) sulphate if carbon electrodes are used instead of those made from copper as referred to in the purification of copper section.
3. Predict the products of the electrolysis of concentrated hydrochloric acid using platinum electrodes.
4. Using your knowledge of electrolysis, predict the likely products of the electrolysis of copper(II) chloride solution, using platinum electrodes. Write electrode equations for the formation of these products.

## Electroplating

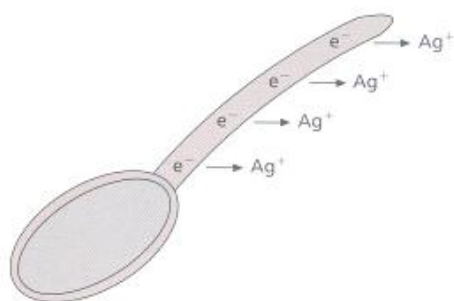
Electroplating is the process involving electrolysis to plate, or coat, one metal with another or a plastic with a metal. Often the purpose of electroplating is to give a protective coating to the metal beneath. For example, bath taps are chromium plated to prevent corrosion, and at the same time are given a shiny, more attractive finish (Figure 6.18).



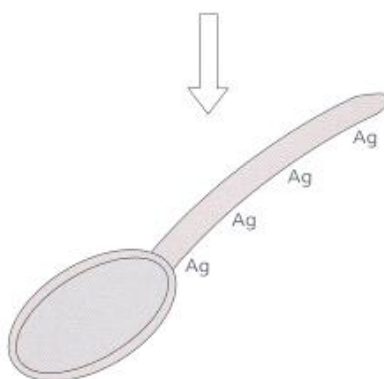
Figure 6.18 This tap has been chromium plated.



a Silver plating a spoon.



The  $Ag^+$  ions are attracted to the cathode, where they gain electrons



A coating of silver forms on the spoon at the cathode

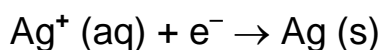
b Explaining silver plating.

Figure 6.19

The electroplating process is carried out in a cell such as the one shown in Figure 6.19a. This is often known as the 'plating bath' and it contains a suitable electrolyte, usually a solution of a metal salt.

For silver plating the electrolyte is a solution of a silver salt. The article to be plated is made the cathode in the cell so that the metal ions move to it when the current is switched on. The cathode reaction in this process is:

silver ions + electrons  $\rightarrow$  silver atoms



### Plating plastics

Nowadays it is not only metals that are electroplated. Plastics have been developed that are able to conduct electricity. For example, the plastic poly (pyrrole) can be electroplated in the same way as the metals we have discussed above (Figure 6.20).



**Figure 6.20** This plastic has been coated with copper by electrolysis.



Figure 6.21 This leaf has been electroplated.

## Questions

1. The leaf in Figure 6.21 has been copper plated. Suggest a suitable method for copper plating the leaf.
2. Explain why copper (II) chloride solution would not be used as an electrolyte in the electrolyte cell used for copper plating.



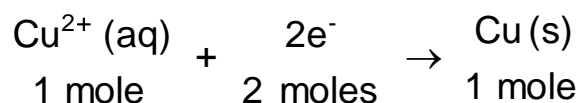
## Calculations in electrolysis

The quantity of electricity flowing through an electrolysis cell is measured in **coulombs** (C). If one **ampere** (A) is passed for one second, the quantity of electricity is said to be 1 coulomb.

$$\begin{array}{ccccc} \text{quantity of electricity} & & \text{current} & & \text{time} \\ \text{(coulombs)} & = & \text{(amperes)} & \times & \text{(seconds)} \end{array}$$

Therefore, if 2 amps flow for 5 seconds then the quantity of electricity passed is 10 coulombs.

In the purification of copper you saw that copper was deposited at the cathode. The electrode equation is:



This equation tells us that 1 mole of copper (II) ions combines with 2 moles of electrons to produce 1 mole of copper metal atoms (63.5 g).

A mole of electrons is called a **faraday**. This unit is named after an English scientist, Michael Faraday (Figure 6.22), who carried out many significant experiments while investigating the nature of magnetism and electricity. So we can say that we need 2 faradays of electricity to form 1 mole of copper atoms (63.5 g) at the cathode during this purification process.

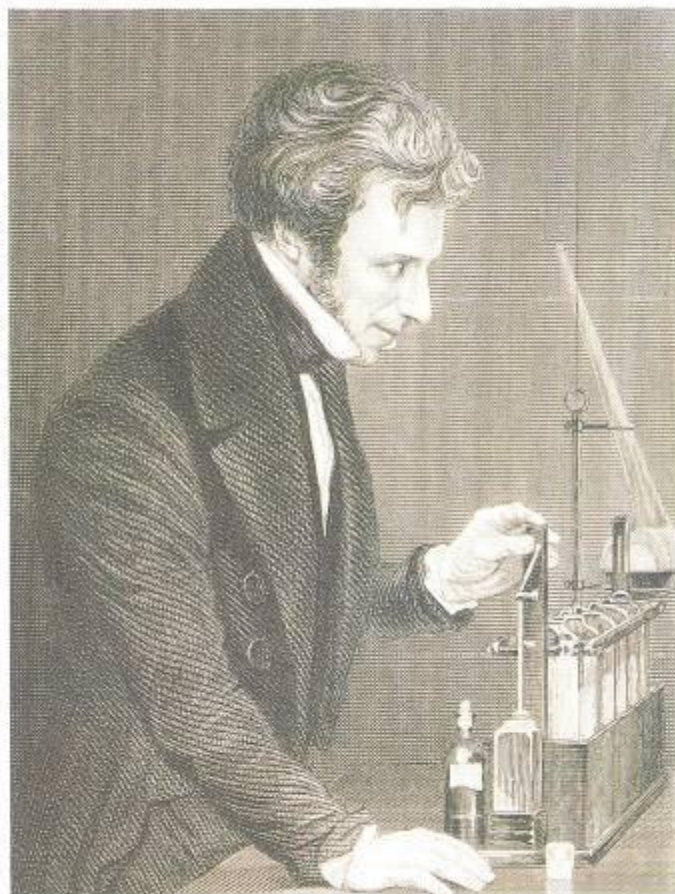


Figure 6.22 In 1883 Michael Faraday was the first scientist to measure the masses of elements produced during electrolysis.

From accurate electrolysis experiments it has been found that:

$$1 \text{ faraday} = 96\,500 \text{ coulombs}$$

Therefore, the quantity of electricity required to deposit 1 mole of copper atoms (63.5 g) is:

$$2 \times 96\,500 = 193\,000 \text{ coulombs (2 faradays)}$$

### Example 1

Calculate the number of faradays required to deposit 10 g of silver on the surface of a fork during an electroplating process. ( $A_r$ : Ag = 108)



1 mole of silver is deposited by 1 faraday. Therefore, 108 g of silver is deposited by 1 faraday. Hence 1 g of silver is deposited by  $\frac{1}{108}$  faraday. Therefore, 10 g of silver are deposited by

$$\frac{1}{108} \times 10 = 0.093 \text{ faradays}$$

### Example 2

Calculate the volume of oxygen gas, measured at room temperature and pressure (rtp), liberated at the anode in the electrolysis of acidified water by 2 faradays. (1 mole of oxygen at rtp occupies a volume of 24 dm<sup>3</sup>.)



1 mole of oxygen gas is liberated by 4 faradays. Therefore, 24 dm<sup>3</sup> of oxygen is liberated by 4 faradays. Hence 12 dm<sup>3</sup> of oxygen would be liberated by 2 faradays.

### Example 3

The industrial production of aluminium uses a current of 25 000 amps. Calculate the time required to produce 10 kg of aluminium from the electrolysis of molten aluminium oxide.



1 mole of aluminium is produced by 3 faradays. 27 g of aluminium is produced by 3 faradays. Hence 1 g of aluminium would be produced by  $\frac{3}{27}$  faradays.

Therefore, 10 000 g of aluminium would be produced by

$$\begin{aligned} \frac{3}{27} \times 10\,000 \text{ faradays} \\ = 1111.1 \text{ faradays} \end{aligned}$$

So, the quantity of electricity

$$= \text{number of faradays} \times 96\,500$$

$$= 1111.1 \times 96\,500$$

$$= 1.07 \times 10^8 \text{ C}$$

Time (seconds)

$$= \frac{\text{coulombs}}{\text{amps}}$$

$$= \frac{1.07 \times 10^8}{25000}$$

$$= 4280 \text{ seconds (71.3 minutes)}$$

## Questions

1. Write equations which represent the discharge at the cathode of the following ions:

a.  $\text{K}^+$

b.  $\text{Pb}^{2+}$

c.  $\text{Al}^{3+}$

and at the anode of:

d.  $\text{Br}^-$

e.  $\text{O}^{2-}$

f.  $\text{F}^-$

2. How many faradays are required to discharge 1 mole of the following ions?



3. How many coulombs are required to discharge 1 mole of the following ions:

(1 faraday = 96 500 coulombs)



4. Calculate the number of faradays required to deposit 6.35 g of copper on a metal surface. ( $A_r$ : Cu = 63.5)

## Checklist

After studying Chapter 6 you should know and understand the following terms.

**Anode** The positive electrode. It is positively charged because electrons are drawn away from it.

**Cathode** The negative electrode. It is negatively charged because an excess of electrons move towards it.

**Coulombs** The unit used to measure the quantity of electricity:

$$1 \text{ coulomb} = 1 \text{ amp} \times 1 \text{ second}$$

**Down's cell** The electrolysis cell in which sodium is extracted from molten sodium chloride to which 60% calcium chloride has been added to reduce the working temperature to 600°C. This cell has a graphite anode surrounded by a cylindrical steel cathode.

**Electrode** A point where the electric current enters and leaves the electrolytic cell. An inert electrode is usually made of platinum or carbon and does not react with the electrolyte or the substances produced at the electrodes themselves.

**Electrolysis** A process in which a chemical reaction is caused by the passage of an electric current.

**Electrolyte** A substance which will carry electric current only when it is molten or dissolved.

**Electroplating** The process of depositing metals from solution in the form of a layer on other surfaces such as metal or plastic.

**Faraday** One mole of electrons, which is equivalent to 96 500 coulombs.

**Hall-Héroult cell** The electrolysis cell in which aluminium is extracted from purified bauxite dissolved in molten cryolite at 900°C. This cell has both a graphite anode and a graphite cathode.

**Membrane cell** An electrolytic cell used for the production of sodium hydroxide, hydrogen and chlorine from brine in which the anode and cathode are separated by a membrane.

**Oxidation** Takes place at the anode and involves a negative ion losing electrons.

**Reduction** Takes place at the cathode and involves a positive ion gaining electrons.



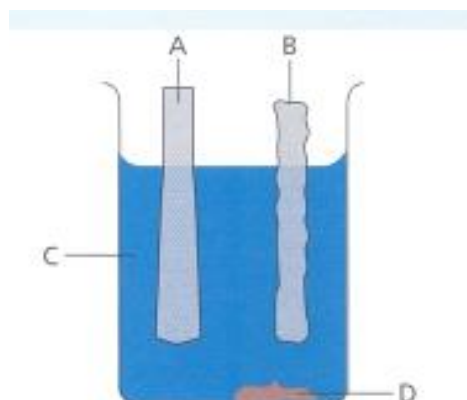
## **Electrolysis and its uses**

### **Additional questions**

1. Explain the meaning of each of the following terms. Use a suitable example, in each case, to help with your explanation.

- a. Anode.
- b. Cathode.
- c. Electrolysis.
- d. Electrolyte.
- e. Anion.
- f. Cation.
- g. Oxidation.
- h. Reduction.

2. Copper is purified by electrolysis, as in the example shown below.



- Name the materials used for the electrodes A and B.
- Name the electrolyte C and substance D.
- Why is substance D of economic importance in respect of this process.
- Write equations for the reactions which take place at the cathode and anode during this process.
- Draw a labelled diagram to show the cell after electrolysis has taken place.
- Why has electrolyte C to be acidified with the dilute sulphuric acid?
- Why does copper have to be 99.99% pure for use in electrical cables?

3. Copy and complete the table below, which shows the results of the electrolysis of four substances using inert electrodes.

<b>Electrolyte</b>	<b>Product at anode (positive electrode)</b>	<b>Product at cathode (negative electrode)</b>
Molten aluminium oxide		Aluminium
Concentrated sodium chloride solution	Chlorine	
Molten lithium chloride		
Silver nitrate solution		Silver

- State what you understand by 'inert electrodes'.
- Explain why the lithium chloride solution becomes progressively more alkaline during electrolysis.
- Explain why solid lithium chloride is a non-conductor of electricity, whereas molten lithium chloride and lithium chloride solution are good conductors of electricity.
- During the electrolysis of molten aluminium chloride ( $\text{AlCl}_3$ ) the carbon anodes are burned away. Explain why this should happen and write balanced chemical equations for the reactions that take place.

4. The industrial production of sodium uses a current of 15 000 amps. Calculate the time required to produce 100 kg of sodium from the electrolysis of molten sodium chloride. (1 faraday = 96500 coulombs.  $A_r$ : Na = 23)

5. Calculate the volume of hydrogen gas, measured at room temperature and pressure, liberated at the cathode in the electrolysis of acidified water by 18 faradays of electricity. (One mole of any gas at rtp occupies a volume of  $24 \text{ dm}^3$ .)

6. Calculate the number of faradays required to produce 10 g of gold deposited on the surface of some jewellery during an electroplating process. (A<sub>r</sub>: Au = 197; note gold in solution is present as the Au<sup>3+</sup> ion)

7. A pupil carried out an experiment in a fume cupboard to find out how electricity affected different substances. Some of the substances were in aqueous solution, others were in the molten state. Carbon electrodes were used in each experiment and she wrote down her results in a table with these headings.

<b>Substance</b>	<b>What was formed at the cathode (–)</b>	<b>What was formed at the anode (+)</b>

Make a table like the one shown and fill it in with what you think happened for each of the substances below.

- a. Molten lead iodide.
- b. Sugar solution.
- c. Silver nitrate solution.
- d. Copper (II) sulphate solution.
- e. Molten sodium bromide.
- f. Ethanol solution.



8. Sodium hydroxide is made by the electrolysis of brine.

a. Draw and label a simplified diagram of the cell used in this process. Make certain that you have labeled on the diagram:

- (i) the electrolyte
- (ii) the material of the electrodes
- (iii) the material of the membrane.

b. Write equations for the reactions which take place at the cathode and anode. State clearly whether a reaction is oxidation or reduction.

c. Give two large-scale uses of the products of this electrolytic process.

d. Comment on the following statement: 'This electrolytic process is a very expensive one'.

e. Both the membrane cell and the older mercury cell make sodium hydroxide of high purity. Explain why the membrane cell is now the preferred way of making sodium hydroxide.